Consider an 0.100 M solution of the weak base ammonia:

$$NH_3 ; K_b = 1.8 \times 10^{-5}$$
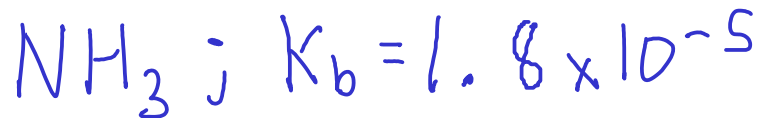
See Appendix I in OpenStax for Kb values

What is the pH?

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^- ; K_b = 1.8 \times 10^{-5} \quad K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

Set up an equilibrium chart:

Species	[Initial]	Δ	[Equilibrium]
NH_4^+	0	+X	X
OH^-	0	+X	X
NH_3	0.100	-X	0.100-X

Let "x" equal the change in ammonium ion concentration

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = \frac{(x)(x)}{(0.100-x)} = 1.8 \times 10^{-5}$$

$$\boxed{\frac{x^2}{0.100-x} = 1.8 \times 10^{-5}}$$

$$\frac{x^2}{0.100 - x} = 1.8 \times 10^{-5}$$

Assume $x \ll 0.100$,
so $0.100 - x \approx 0.100$

$$\frac{x^2}{0.100} = 1.8 \times 10^{-5}$$

$$x = 0.00134164\,08 = [OH^-]$$

So, $pOH = -\log(0.00134164\,08) = 2.87$

$pH + 2.87 = 14.00$

$$pH = \boxed{11.13}$$

$pOH = -\log[OH^-]$

$pH + pOH = 14.00$

If we solve this problem using the quadratic equation instead of our simplifying assumption, we get a pH of 11.13 ... same as we got here.

Compare pH to the pH of an 0.100 M solution of the strong base NaOH:

$$pH_{,NH_3} = 11.13$$

$$NaOH \longrightarrow Na^+ + OH^-$$

$$[OH^-] = [NaOH]_{nominal} = 0.100\,M$$
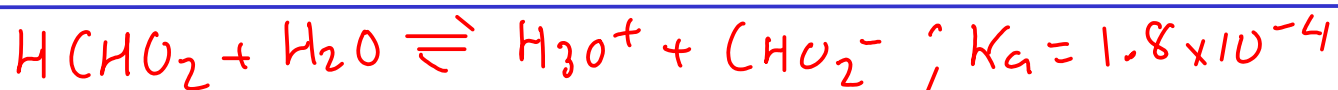
$$pOH = 1.00$$

$$pH = 13.00$$

The stronger the base:

- the higher the pH will be for a solution of given concentration
- the higher the HYDROXIDE concentration (compared to the nominal base concentration)

Find the pH and the degree of ionization for an 0.10 M solution of formic acid: $HCHO_2$

$$HCHO_2 + H_2O \rightleftharpoons H_3O^+ + CHO_2^- \quad ; K_a = 1.8 \times 10^{-4}$$

$$K_a = \frac{[H_3O^+][CHO_2^-]}{[HCHO_2]} = 1.8 \times 10^{-4}$$

Species	[Initial]	Δ	[Equilibrium]
H_3O^+	0	$+x$	x
CHO_2^-	0	$+x$	x
$HCHO_2$	0.10	$-x$	$0.10 - x$

Let "x" equal the change in hydronium ion concentration

$$\frac{(x)(x)}{(0.10 - x)} = 1.8 \times 10^{-4}$$

$\downarrow$ Assume $x << 0.10$

$$\frac{x^2}{0.10} = 1.8 \times 10^{-4}$$

$x = 0.004242640 7 = [H_3O^+]$

So, pH $= 2.37$

... but what about degree of ionization?

$$HCHO_2 + H_2O \rightleftharpoons H_3O^+ + CHO_2^-$$

Degree of Ionization (DOI) is the fraction of a weak acid or base that ionizes in solution. Here, that would be:

$$\frac{[CHO_2^-]}{[HCHO_2]_{nominal}} = \frac{[H_3O^+]}{[HCHO_2]_{nominal}}$$

$$\frac{0.0042426407 \, m \, H_3O^+}{0.10 \, m \, HCHO_2} = \boxed{0.042 = DOI \text{ for this } 0.10 \, m \, HCHO_2}$$

Often, we express this as a percentage ... called PERCENT IONIZATION

$$\% \text{ ionization} = DOI \times 100 = 4.2\% \text{ ionization}$$

When you do Experiment 16A. By Le Chateleir's Principle, adding water to the equilibrium should force it to the right - meaning that more acid will ionize - even as the pH goes up!. Therefore, the degree of (or percent) ionization should INCREASES as the concentration of the acid DECREASES. Check this with your experiment 16A data on acetic acid.

An aqueous solution of 0.25 M trimethylamine has a pH of 11.63. What's the experimental value of Kb?

$$(CH_3)_3N$$
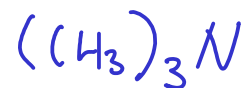
$$(CH_3)_3N + H_2O \rightleftharpoons (CH_3)_3NH^+ + OH^-$$

$$K_b = \frac{[(CH_3)_3NH^+][OH^-]}{[(CH_3)_3N]}$$

We don't know Kb, so how do we proceed? Let's make an equilibrium chart to try to reduce the number of variables!

Species	[Initial]	Δ	[Equilibrium]
$(CH_3)_3NH^+$	0	+x	x
OH^-	0	+x	x
$(CH_3)_3N$	0.25	-x	0.25 - x

Let "x" equal the change in trimethylammonium ion concetration.

Plug into the Kb expression

$$\frac{(x)(x)}{(0.25 - x)} = K_b$$

$$\frac{x^2}{0.25 - x} = K_b$$

Now what? We have fewer variables now, but there are still two...

$$\frac{x^2}{0.25 - x} = K_b$$

If we had some other way to find the value of "x", we could then get Kb.
Use the pH to find the pOH, and then (OH-) ... which equals "x".

$$11.63 + pOH = 14.00 \qquad pH + pOH = 14.00$$

$$pOH = 2.37$$

$$[OH] = 0.004265795 2 \qquad [OH^-] = 10^{-pOH}$$

$$So, \; x = 0.0042657952$$

$$K_b = \frac{x^2}{0.25 - x} = \frac{(0.0042657952)^2}{0.25 - 0.0042657952} = \boxed{7.4 \times 10^{-5} = K_b}$$